Chapter 19 ATOMIC SPECTRA

- The word atom is a Greek word meaning indivisible, i.e.… 'An ultimate particle which cannot be further subdivided'.
- The idea that all matter ultimately consists of extremely small particles was conceived by ancient Indian and Greek philosophers.
- The old concept was put on firm footing by John Dalton in the form of atomic theory which he developed in the years 1803-1808.
- The theory was a landmark in the history of chemistry. According to this theory, atom is the smallest indivisible part of matter which takes part in chemical reactions.
- Atom is neither created nor destroyed.
- Atoms of the same element are similar in size, mass and characteristics; however, atoms of different elements have different size, mass and characteristics.
- In 1833, Michael Faraday showed that there is a relationship between matter and electricity.
- This was the first major breakthrough to suggest that atom was not a simple indivisible particle of all matter but was made up of smaller particles.
- Discovery of electrons, protons and neutrons discarded the indivisible nature of the atom proposed by John Dalton.
- The complexity of the atom was further revealed when the following discoveries were made in subsequent years:
  - ✓ Discovery of cathode rays.
  - ✓ Discovery of positive rays.
  - ✓ Discovery of X-rays.
  - ✓ Discovery of radioactivity.
  - ✓ Discovery of isotopes and isobars.
  - ✓ Discovery of quarks and the new atomic model.
- During the past 100 years, scientists have made contributions which helped in the development of modern theory of atomic structure.
- The works of **J.J Thomson** and **Ernest Rutherford** actually laid the foundation of the modern picture of the atom.
- It is now believed that the atom consists of several particles called subatomic particles like electron, proton, neutron, positron, neutrino, meson, etc. Out of these particles, the electron, the proton and the neutron are called fundamental particles and are the building blocks of the atom.

Spectrum

- A band of radiation obtained by the dispersion of radiation is called spectrum.
- Spectrum is the impression produced on a screen when radiations of particular wavelengths are analysed through a prism or diffraction grating.
- A prism is made up of calcium carbonate.

Spectrum are of two types:

1) Line spectrum;
2) Continuous spectrum

Line spectrum

- The spectra where wavelength is separated by different boundaries/Lines.
- It is also known as atomic spectra.
- In this chapter mainly we will study Atomic spectra.

- ⇒ Electrons on excitation absorbs energy and on de-excitation emit energy of specific wavelength.
- ⇒ We know that in atom there are different shells, so the cause of line spectra is excitation and de-excitation of electron between shells.
- ⇒ There are two types of line spectra:

1) Line absorption spectrum
2) Line emission spectrum

Line emission spectrum	Line absorption spectrum
It gives bright lines (colored) on the dark background.	It gives dark lines on the bright background.
Radiations from emitting source are analyzed by the spectroscope.	It is observed when the white light is passed through the substance and the transmitted radiations are analyzed by the spectroscope.
It may be continuous (if source emits white light) and may be discontinuous (if the source emits colored light).	These are always discontinuous.

1) **Line absorption spectra:**

- ✓ It is the electromagnetic spectrum, broken by a specific pattern of dark lines or bands, observed when radiation traverses a particular absorbing medium and through a spectroscope.
- ✓ The absorption pattern of an element is unique and can be used to identify the substance of the medium.
- ✓ When white light is passed through a gas, the gas is absorbing light of certain wavelength.
- ✓ The bright background on the photographic plate is then crossed by dark lines that corresponds o those wavelengths which are absorbed by the gas atoms, resulting in transition of an atom from lower energy states to higher energy states.

The spectrum of sunlight has dark lines called Fraunhofer lines. These lines are produced when the light emanating from the core of the sun passes through the layer of cooler gas. This layer absorbs light of certain wavelengths corresponding to the elements present in the cooler gas. This results in dark lines (absorption of certain wavelengths) on a brighter background. Fraunhofer lines reveal the composition of the star.

2) **Line emission spectrum:**

- ✓ When an electric current pass through a gas which is at less than atmospheric pressure, it gives energy to the gas.
- ✓ This energy is then given out as light of several definite wavelengths (colors). This is called a emission line spectrum.
- ✓ These are caused when an electron hops from excited states to lower states.
- ✓ Different The wavelength of emission lines of different elements has emissions of different wavelengths.
- ✓ For one element the emission spectrum is unique for each element.

Emission spectra are of two types:

(a) Continuous spectra and (b) Discontinuous spectra.

(a) Continuous spectra:

- ✓ When white light is allowed to pass through a prism, it gets resolved into several colors.
- ✓ The spectrum is a rainbow encolours, i.e., violet merges into blue, blue into green, and so on.
- ✓ This is a continuous spectrum.

(b) Discontinuous spectra:

- ✓ When gases or vapors of a chemical substance are heated in an electric arc or in a Bunsen flame, light is emitted.
- ✓ If a ray of this light is passed through a prism, a line spectrum is produced.
- ✓ This spectrum Consists of a limited number of lines, each of which corresponds to a different wavelength of light.
- ✓ The line spectrum of each element is unique.
- ✓ Hydrogen spectrum is an example of line emission spectrum or atomic emission spectrum.
- ✓ When an electric discharge is passed through hydrogen gas at low pressure, a bluish light is emitted.
- ✓ When a ray of this light is passed through a prism, a discontinuous line spectrum of several isolated sharp lines is obtained.

- ✓ The wavelengths of various lines show that these lines lie in visible, ultraviolet and infrared regions.
- ✓ All these lines observed in the hydrogen spectrum can be classified into six series

Spectral series	Discovered by	Appearing in
Lyman series	Lyman	Ultra violet
Balmer series	Balmer	Visible
Paschen series	Paschen	Near Infra-red
Bracket series	Bracket	Mid infra-red
Pfund series	Pfund	Far infra-red

Concept

The only electron in the hydrogen atom resides under ordinary conditions on the first orbit. When energy is supplied the electron moves to higher energy shells depending on the amount of energy absorbed.
When this electron returns to any of the lower energy shells, it emits energy. Lyman series is formed when the electron returns to the lowest energy state while Balmer series is formed when the electron returns to second energy shell. Similarly, Paschen, Brackett and Pfund series are formed when electron returns to the third, fourth and fifth energy shells from higher energy shells respectively.

Ritz Equation

- ✓ He presented a mathematical formula to find the wavelengths of various hydrogen lines.;

$$\bar{v} = \frac{1}{\lambda} = \frac{v}{c} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

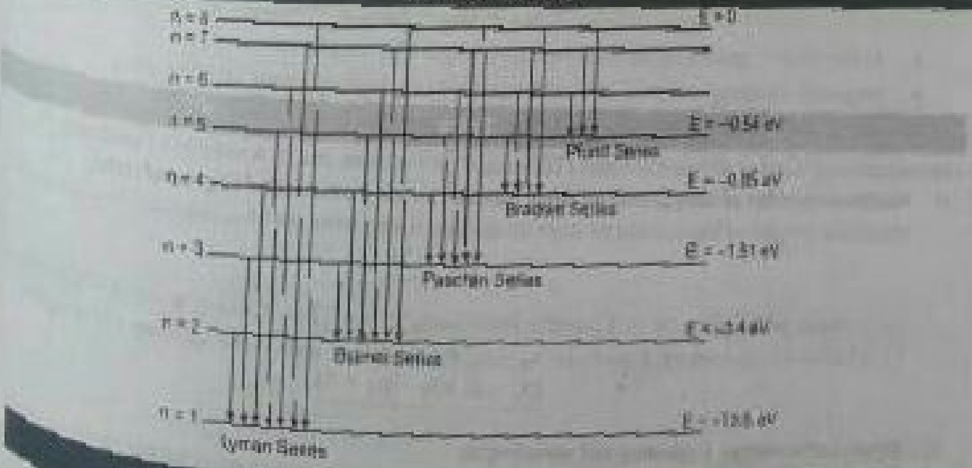
Where, R is a universal constant, known as Rydberg constant. Its value is 109,678 cm^{-1}, n_1 and n_2 are integers (such that $n_2 > n_1$). For a given spectral series, n_1 remains constant while n_2 varies from line to line in the same series.

- ✓ The value of n_1 =1, 2, 3, 4 and 5 for the Lyman, Balmer, Paschen, Brackett and Pfund series respectively. n_2 is greater than n_1 by at least 1

Values of n_1 and n_2 for various series

Spectral Series	Value of n_1 (Ground state)	Value of n_2 (Excited state)
Lyman series	1	2,3,4,5,...
Balmer series	2	3,4,5,6,...
Paschen series	3	4,5,6,7,...
Bracket series	4	5,6,7,8,...
Pfund series	5	6,7,8,9,...

Spectral series;



Lyman series;

$n_1 = 1$ and $n_2 = 2,3,4, \ldots \infty$

The wavelength emitted is given by:

$$As, f = \frac{c}{\lambda} = \frac{\Delta E}{h} \quad or\ \lambda = \frac{hc}{\Delta E}$$

Since h and c are constants, ΔE corresponds to definite energy; thus, each transition from one energy level to another will produce a light of definite wavelength. This is actually observed as a line in the spectrum of hydrogen atom.

3) **Number of Absorption Spectral Lines:**
At ordinary temperatures almost all the atoms remain in their lowest energy level ($n = 1$) and, therefore Absorption transition can start only from the lowest energy level i.e., $n = 1$ level (not from $n = 2, 3, 4,$ ..... levels). Hence, only Lyman series is found in the absorption spectrum of hydrogen atom (which as in the emission spectrum, all the series are found)

$$Number\ of\ absorption\ spectral\ lines = (n - 1)$$

4) **For maximum wavelength:**
First line of the series is called first member, for this line wavelength is maximum (λ_{max})
For maximum wavelength $if\ n_1 = n\ then\ n_2 = n + 1$
So.

$$\lambda_{max} = \frac{n^2(n+1)^2}{(2n+1)R}$$

5) **For minimum wavelength:**
Last line of the series is called series limit, for this line wavelength is minimum (λ_{min})
For minimum wavelength $if\ n_2 = \infty, n_1 = n$
So:

$$\lambda_{min} = \frac{n^2}{R}$$

6) The ratio of first member and series limit can be calculated as:

$$\frac{\lambda_{max}}{\lambda_{min}} = \frac{(n+1)^2}{(2n+1)}$$

Dalton's Atomic Model

- All matter is made of tiny particles called atoms. Atoms are indivisible and indestructible.
- All atoms of a given element are identical in mass and properties, while atoms of different elements differ in mass and properties.
- All matter is made up of hydrogen atoms. The mass and radius of heaviest atom is about 250X and 10X of the than that of the hydrogen atom, respectively.
- Atoms are stable and electrically neutral.

Reason of Failure of model: The discovery of electron by J.J. Thomson (1897) proved that atoms are not indivisible hence, the model is no longer valid.

Thomson's Atomic Model (or Plum Pudding Model)

- In this model, the atom is composed of electrons (which Thomson still called "corpuscles") surrounded by a soup of positive charge to balance the electrons' negative charges, like negatively charged "raisins" surrounded by positively charged "pudding".
- *Achievements of model:* Explained successfully the phenomenon of thermionic emission, photoelectric emission, and ionization.
- *Failure of the model:*
(a) It could not explain the line spectrum of H-atom.
(b) It could not explain the Rutherford's α ~ particle scattering (Rutherford gold foil) experiment.

Rutherford's Experiment and Atomic Model

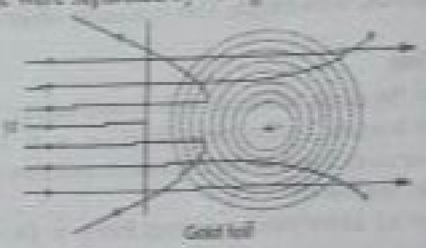
- This experiment is also called α ~ *scattering experiment*.

Results of Experiment:

- It was seen that in the experiment that when the α - particles were fired at the gold foil, some of the particles (<1 in 8000) bounced off the metal foil in all directions, some right back at the source. This should have been impossible according to Thomson's model; the alpha particles should have all gone straight through.
Obviously, those particles had encountered an electrostatic force far greater than Thomson's model suggested they would, which in turn implied that the atom's positive charge was concentrated in a much

tinier volume than Thomson imagined. This was possible only in the case when there exists a solid positive mass confining in a very narrow space.

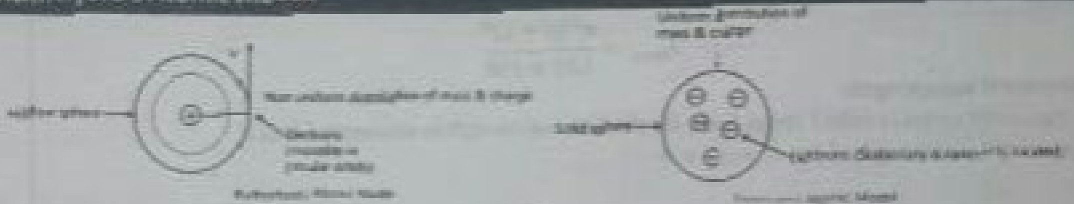
(b) However, most of the α - particles just flew straight through the foil. This suggested that those tiny spheres of intense positive charge were separated by vast gulfs of empty space.



(c) $N \propto (Nuclear\ charge)^2$

where N represents → No of particles scattered per unit time

⇒ Rutherford's Atomic Model:





✓ The whole positive charge and almost whole mass of an atom (leaving aside the mass of revolving e^- in various circular orbits) remains concentrated in nucleus of radius of the order of 10^{-8} m.

✓ $\Sigma + q(+)$ve on proton in a nucleus = $\Sigma - q(-)$ve on e^- in various circular orbits & hence, the atom is electrically neutral.

✓ The necessary centripetal force for revolving round the nucleus in circular orbit is provided by coulomb's electrostatic force of attraction

⇒ Reason of failure of model:

(a) It could not explain the line spectrum of H-atom.

Justification: Asper Maxwell's electromagnetic theory, every accelerated moving charged particle emits energy in the form of electromagnetic waves and therefore, the frequency of a Modern Physics moving in a circular orbit around the nucleus will steadily decline, resulting in the continuous emission of lines thereby mandating that the spectrum of an atom be continuous, but in reality, one obtains line spectrum for atoms.



Motion of electron in atom

(b) It could not explain the stability of atoms.

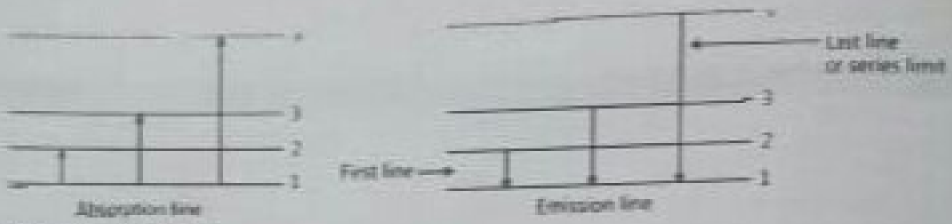
Justification: Since revolving electron continuously radiates energy, the radii of circular path will continuously decrease and in a time of about 10-8 s the revolving electron must fall down in a nucleus by adopting a spiral path.

Bohr's Model of Hydrogen atom

⇒ Bohr combined the concepts of classical physics with quantum mechanics to propose his model for H or H-like atoms. This model is based on law of conservation of angular momentum.

⇒ This model is also known as "quantum mechanical model of the atom" because This model was based on the quantum theory.

⇒ of radiation and the classical laws of physics.



Maximum wavelength:

For 1^{st} line or series beginning $n_1 = 1,\ n_2 = 2;\ \frac{1}{\lambda} = R\left[\frac{1}{1^2} - \frac{1}{2^2}\right];\ \lambda_{max} = \frac{4}{3R}$

Minimum wavelength:

For series limit or last line $n_1 = 1,\ n_2 = \infty;\ \frac{1}{\lambda} = R\left[\frac{1}{1^2} - \frac{1}{\infty^2}\right];\ \lambda_{min} = \frac{1}{R}$

Spectral series	Transition	λ_{max}	λ_{min}	$\frac{\lambda_{max}}{\lambda_{min}}$	Region	Description
Lyman series	$n_2 = 2, 3, 4 \ldots \infty$ $n_1 = 1$	$\frac{4}{3R}$	$\frac{1}{R}$	$\frac{4}{3}$	U-v	-
Balmer series	$n_2 = 3, 4, 5 \ldots \infty$ $n_1 = 2$	$\frac{36}{5R}$	$\frac{4}{R}$	$\frac{9}{5}$	Visible	Balmer series is found only in emission spectrum.
Paschen series	$n_2 = 4, 5, 6 \ldots \infty$ $n_1 = 3$	$\frac{144}{7R}$	$\frac{9}{R}$	$\frac{16}{7}$	IR	Paschen series is also found only in emission spectrum.
Bracket series	$n_2 = 5, 6, 7 \ldots \infty$ $n_1 = 4$	$\frac{400}{9R}$	$\frac{16}{R}$	$\frac{25}{9}$	IR	Brackett series is also found only in emission spectrum.
Pfund series	$n_2 = 6, 7, 8 \ldots \infty$ $n_1 = 5$	$\frac{900}{11R}$	$\frac{25}{4}$	$\frac{36}{11}$	IR	Pfund series is also obtained only in emission spectrum.

Some Important Points

- The minimum wavelength of a series (Lyman, Balmer, Paschen, Brackett etc.) correlates with the ionization potential of the electron from that shell.
- Wavelength of first line is maximum and last line is minimum

$$\lambda_{pf} > \lambda_{Br} > \lambda_{pas} > \lambda_{Balmer} > \lambda_{Lyman}$$

- As the order of spectral series increases, wavelength also usually increases.
- Frequency of energy emission in Lyman transitions are highest among all other series.

Trick

Determination of Number of Spectral Lines (Theoretical) in Emission and in Absorption Transitions

i) ***Maximum number of lines produced when electron jumps from excited to ground state:***
Maximum number of lines produced when an electron jumps from nth level to ground level is equal to

$$\frac{n(n-1)}{2}$$

For example: in the case of $n = 4$, number lines produced is 6. ($4 \to 3, 4 \to 2, 4 \to 1, 3 \to 2, 3 \to 1, 2 \to 1$) When an electron returns from n_2 to n_1 state, the number of lines in the spectrum will be equal to

$$\frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

2) ***Difference in energy, frequency and wavelength:***
If the electron comes back from energy level having energy E_2 to energy level having energy E_1, then the difference may be expressed in terms of energy of photon as:

$$E_2 - E_1 = \Delta E = h\nu$$

The frequency of emitted radiation is given by:

$$f = \frac{\Delta E}{h}$$

1) The atom has a nucleus where all the protons and neutrons are present. The size of the nucleus is very small. It is present at the centre of the atom.
2) Negatively charged electrons are revolving around the nucleus in the same way as the planets are revolving around the sun. The path of the electron is circular. The force of attraction between the nucleus and the electron is equal to centrifugal force of the moving electron.
3) Out of infinite number of possible circular orbits around the nucleus, the electron can revolve only on those orbits whose angular momentum is an integral multiple of

$$or\ 2\pi r = n\lambda\ (constructive\ interference)$$

$$or\ 2\pi r = \frac{nh}{mv} \quad or\ mvr = \frac{nh}{2\pi}$$

The angular momentum can have values such as, $\frac{h}{2\pi}, \frac{2h}{2\pi}, \frac{3h}{2\pi}$ etc

So, in an orbit, waves are always formed in whole numbers, it cannot have a fractional value. Thus, the angular momentum is quantized. The specified or circular orbits (quantized) are called stationary orbits.

4) By the time, the electron remains in anyone of the stationary orbits, it does not lose energy. Such a state is called ground or normal state.

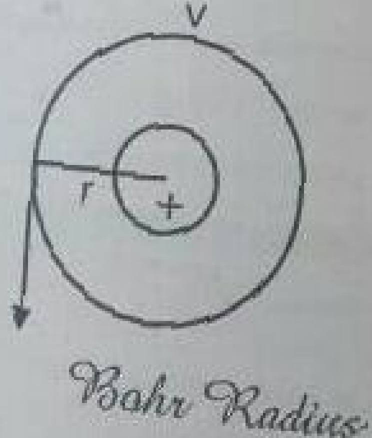
In the ground state, potential energy of electron will be minimum, hence it will be the most stable state.

5) Each stationary orbit is associated with a definite amount of energy. The greater is the distance of the orbit from the nucleus, more shall be the energy associated with it. These orbits are also called energy levels and are numbered as 1, 2, 3, 4, ... or K, L, M, N, ... from nucleus outwards;

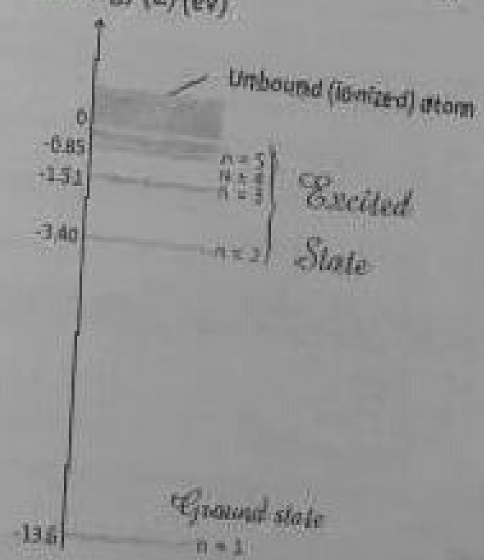
$$E_1 < E_2 < E_3 < E_4 ...$$

$$OR \quad (E_2 - E_1) > (E_3 - E_2) > (E_4 - E_3) ...$$

6) According to Bohr, if an electron revolves in these orbits the electron neither radiates nor absorbs any energy.



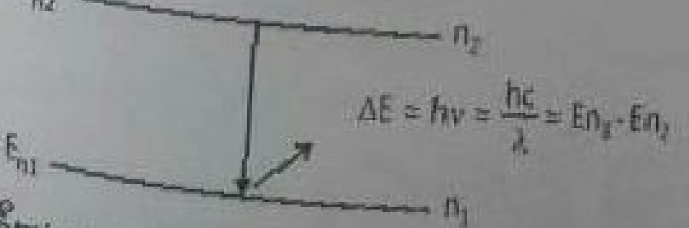
Bohr Radius



Energy Level Diagram

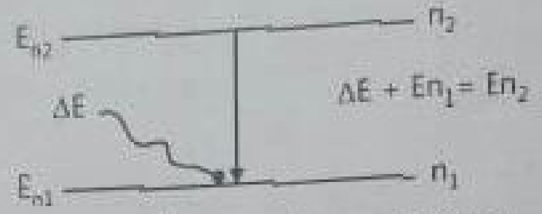
7) Emission of energy

Where n = principle quantum no.
E_n = energy of e^- in nth orbit



Emission of energy by an electron

8)



Absorption of energy by an electron

9) When the electron moves from inner to outer orbit by absorbing definite amount of energy, the new state of the electron is said to be excited state.

Using the above postulates, Bohr calculated the radii of various stationary orbits, the energy associated with each orbit and explained the spectrum of hydrogen atom:

1) Determination of radius of circular path (orbit):

Consider an electron of mass 'm' and charge 'q' revolving around the nucleus of charge Ze (Z =atomic number). Let 'v' be the tangential velocity of the revolving electron and 'r' the radius of the orbit;

$$r = \frac{n^2h^2}{4\pi^2mkZe^2}$$

- ✓ Greater is the value of 'n', larger is the size of atom.
- ✓ On the other hand, greater is the value of 'Z', smaller is the size of the atom.
- ✓ Across a period from left to right, atomic number 'Z' increases with constant value of 'n' hence atomic radius decreases towards right.
- ✓ On moving down the group, both 'Z' and 'n' increase but due to shielding, Z (effective nuclear charge) remains same. Hence, on moving downwards, atomic radius increases due to increase in 'n'.
- ✓ As for radius of 1st orbit n=1, so putting n=1 in equation above, and putting the values of h, m, e, k we get;

$$0.529 \times 10^{-8} \times 1^2 = 0.529 \times 10^{-8}cm = 0.529 \times 10^{-10}m$$

- ✓ Radius of 2nd orbit:

$$0.529 \times 10^{-8} \times 2^2 = 2.11 \times 10^{-8}cm = 2.11 \times 10^{-10}m$$

- ✓ Radius of 3rd orbit:

$$0.529 \times 10^{-8} \times 3^2 = 4.76 \times 10^{-8}cm = 4.76 \times 10^{-10}m$$

- ✓ And so on:

$$r_n = r_1 \times n^2 \quad for\ hydrogen\ atom$$

- ✓ And; $r_n = 0.529 \times \frac{n^2}{Z}\ A^{\circ}$ $\quad$ *(For hydrogen like species)*

MCQ: What will be the ratio of the area of circular orbits in doubly ionized lithium (Li) atom in 2nd & 3rd Bohr's orbit?

Solution: According to Bohr's theory as: $r_n \propto n^2$ but $A \propto r^2$, therefore $A \propto n^4$

$$\Rightarrow \frac{A_2}{A_3} = \frac{(2)^4}{(3)^4} = \frac{16}{81}$$

2) Energy of an electron

Let the total energy of the electron be E. It is the sum of kinetic energy and potential energy.

- ✓ The total energy of an electron is given by:

$$E = -\frac{2\pi^2Z^2k^2me^4}{n^2h^2}$$

- ✓ As for hydrogen atom Z=1 so the energy equation for hydrogen atom is;

$$E = -\frac{2\pi^2k^2me^4}{n^2h^2}$$

- ✓ By putting values, we get:

$$E = -\frac{21.79 \times 10^{-19}}{n^2} = -\frac{R_H}{n^2} \quad (Where\ R_H = 21.79 \times 10^{-19})$$

- ✓ $E_n = -\frac{13.6}{n^2}$ *eV per atom.*

energy and potential energy of an electron increases, whereas the kinetic energy decreases.

⇨ With the increase in principal quantum number, the difference between any two consecutive energy level decreases.

Success of Bohr's theory

1) Bohr successfully combined Rutherford's model with the Planck hypothesis on the quantified energy states at atomic level.
2) Bohr's theory explained the atomic emission and absorption spectra.
3) It explained the general characteristics of the periodic table.
4) Bohr's theory offered the first "working" model for the atom.

Short Coming of Bohr's Model

1) Bohr's model holds true only for atoms with one electron. E.g. $H, He^{+}, Li^{+2}, Na^{+2}$.
2) Bohr's model posits circular orbits whereas according to Sommerfeld these are elliptical.
3) The model could not explain the intensity of spectral lines.
4) It assumes the nucleus to be stationary, but it also rotates on its own axis.
5) It failed to account for the minute structure in spectrum line.
6) The model offered no explanation for the Zeeman effect (spliting up of spectral lines in magnetic field) and Stark effect (splitting up of spectral lines in electric field)
7) Doublets observed in the spectrum of some of the atoms like sodium (5890 Å & 5896 Å) could not be explained by Bohr's model.

Failure of Bohr's Model

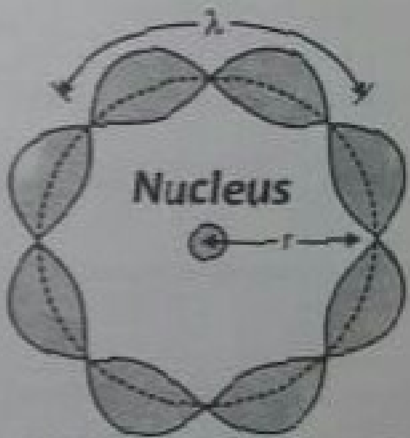
1) The Bohr model is applicable to hydrogenic atoms with a single electron. All attempts to use Bohr's Model to analyze atoms with more than one electron failed as Bohr's model deals only with interaction between the electron and the positively charged nucleus but does not account for the interaction of an electron with other electrons as would be the case with multi-electron atoms.
2) While the Bohr's model correctly predicts the frequencies of the light emitted by hydrogenic atoms, it cannot predict the relative intensities of spectral lines. Some frequencies in the hydrogen emission spectrum, for example, have weak intensity while others have strong intensity. Bohr's model is unable to account for the intensity variations.

De-Broglie waves and the Hydrogen atom

⇒ WHY?

In Bohr's model of the atom, it is stated that the angular momentum of the electron orbiting around the nucleus is quantized (that is $L_n = \frac{nh}{2\pi}$; n = 1,2,3 - - -) Why is it that the values of angular momentum are only integral multiples of $\frac{h}{2\pi}$. This postulate can be derived with the help of de Broglie concept of wave nature of electron.

⇒ Consider an electron moving in a circular orbit around nucleus. The wave train would be associated with the circular orbit as shown in figure. If the two ends of an electron wave meet to give a regular series of crests and troughs, the 'electron wave is said to be in phase, i.e., the circumference of Bohr orbit is equal to whole number multiple of the wavelength of the electron wave;



[illegible]

So,

$$2\pi r = n\lambda \Rightarrow \lambda = \frac{2\pi r}{n} \quad ...(i)$$

From the De-Broglie equation;

$$\lambda = \frac{h}{mv} \quad ...(ii)$$

Thus;

$$\frac{h}{mv} = \frac{2\pi r}{n} \Rightarrow mvr = n.\frac{h}{2\pi}$$

and so Angular momentum $= n.\frac{h}{2\pi}$

- ✓ $E_n = -\frac{313.6}{n^2}$ Kcal/mol
- ✓ $E_n = -\frac{1312}{n^2}$ kJ/mol
- ✓ if r ↑ the TE ↑
- ✓ if r = ∞ then TE = 0
- ✓ $TE = -KE = \frac{PE}{2}$ at any H like atom
- ✓ Kinetic energy in n^{th} shell $= \frac{13.6 \times Z^2}{n^2}$ eV and Kinetic energy is a positive quantity.

$$r \uparrow, KE \downarrow$$
$$\text{when } r = \infty \text{ then } KE = 0$$

- ✓ Potential energy in n^{th} shell $= \frac{-27.2 \times Z^2}{n^2}$ eV

$$\text{Potential energy of an electron is negative}$$
$$r \uparrow, PE \uparrow$$
$$\text{When } r = \infty \text{ then } PE = 0$$

Substituting the values of n: 1, 2, 3, 4, ..., etc., the energy of electron in various energy shells in hydrogen atom can be calculated.

Energy Shell	E (Joule per atom)	E (eV per atom)	E (kcal per mol)
1	-21.79×10^{-19}	−13.6	−313.6
2	-5.44×10^{-19}	−3.4	−78.4
3	-2.42×10^{-19}	−1.51	−34.84
4	-1.36×10^{-19}	−0.85	−19.6
–	–	–	–
–	–	–	–
∞	0	0	0

- ✓ $E_n = \frac{E_1}{n^2}$ (For hydrogen atom)
- ✓ and $E_n = E_1 \times \frac{Z^2}{n^2}$ (For hydrogen like species like He^+, Li^{2+} etc)
- ✓ Since, n can have only integral values, it follows that total energy of the electron is quantized.
- ✓ The negative sign indicates that the electron is under attraction towards nucleus, i.e., it is bound to the nucleus.
- ✓ The electron has minimum energy in the first orbit and its energy increases as n increases, i.e., it becomes less negative.
- ✓ The electron can have a maximum energy value of zero when n = ∞. The zero energy means that the electron is no longer bound to the nucleus, i.e., it is not under attraction towards nucleus.

Orbital velocity

- ✓ The velocity of electron in an orbit is given by:

$$v = \frac{Z}{n}\left(\frac{2\pi e^2}{h}\right) \Rightarrow v = \frac{Z}{n} \times 2.188 \times 10^8 \ cm/sec$$

- ✓ For hydrogen Z = 1 so:

$$v = \frac{2.188 \times 10^8}{1} = 2.188 \times 10^8 \ cm/sec$$

- ✓ The velocity of electron is 2^{nd} orbit is given by:

$$v_2 = \frac{1}{2} \times 2.188 \times 10^8$$

- ✓ The velocity of electron in 3^{rd} orbit is given by:

$$v_3 = \frac{1}{3} \times 2.188 \times 10^8$$

- ✓ Here is the velocity of electron in 1^{st}, 2^{nd} and 3^{rd} orbits.
- ✓ The ratio of velocity in 1^{st} orbit to velocity in 2^{nd} orbit:

$$\frac{v_1}{v_2} = \frac{2}{1}$$

- ✓ The ratio of velocity in 3^{rd} orbit to velocity in 1^{st} orbit:

$$\frac{v_3}{v_1} = \frac{1}{3} \quad \text{and so on} - - -$$

Some important Points Regarding energy:

- ✓ With the increase in principal quantum number n (relative overall energy of each orbital), both total

This proves that the de Broglie and Bohr concepts are in perfect agreement with each other.

Summary

Physical Quantity	Formula	Ratio formulae of hydrogen atom
Radius of Bohr orbit	$r_n = \frac{n^2h^2}{4\pi^2 m^2 kZe^2}$; $r_n = \frac{0.53n^2}{Z} A^0$	$r_1:r_2:r_3 \dots r_n = 1:4:9 \dots n^2$
Velocity of electron in nth orbit	$v_n = \frac{2\pi kZe^2}{nh}$; $v_n = 2.2\ 10^6 \frac{Z}{n}$	$v_1:v_2:v_3 \dots v_n = 1:\frac{1}{2}:\frac{1}{3} \dots \frac{1}{n}$
Momentum of electron in orbit	$P_n = \frac{2\pi mke^2Z}{nh}$; $P_n \propto \frac{Z}{n}$	$P_1:P_2:P_3 \dots P_n = 1:\frac{1}{2}:\frac{1}{3} \dots \frac{1}{n}$
Angular velocity of an electron	$\omega_n = \frac{8\pi^3k^2Z^2me^4}{n^3h^3}$; $\omega_n \propto \frac{Z^2}{n^3}$	$\omega_1:\omega_2:\omega_3 \dots \omega_n = 1:\frac{1}{8}:\frac{1}{27} \dots \frac{1}{n^3}$
Time period of an electron	$T_n = \frac{n^3h^3}{4\pi^2k^2Z^2me^4}$; $T_n \propto \frac{n^3}{Z^2}$	$T_1:T_2:T_3 \dots T_n = 1:8:27 \dots n^3$
Frequency of an electron	$f_n = \frac{4\pi^2k^2Z^2me^4}{n^3h^3}$; $f_n \propto \frac{Z^2}{n^3}$	$f_1:f_2:f_3 \dots f_n = 1:\frac{1}{8}:\frac{1}{27} \dots \frac{1}{n^3}$
Angular momentum	$L_n = \frac{nh}{2\pi}$; $L_n \propto n$	$L_1:L_2:L_3 \dots L_n = 1:2:3 \dots n$
Centripetal acceleration	$a_n = \frac{16\pi^4k^3Z^3me^6}{n^4h^4}$; $a_n \propto \frac{Z^3}{n^4}$	$a_1:a_2:a_3 \dots a_n = 1:\frac{1}{16}:\frac{1}{81} \dots \frac{1}{n^4}$
Kinetic energy	$KE = \frac{13.6 \times Z^2}{n^2} eV$; $K.E \propto \frac{Z^2}{n^2}$	$E_{KE1}:E_{KE2}:E_{KE3} \dots E_{KEn} = 1:\frac{1}{4}:\frac{1}{9} \dots \frac{1}{n^2}$
Potential energy	$PE = \frac{-27.2 \times Z^2}{n^2} eV$; $PE \propto \frac{Z^2}{n^2}$	$E_{PE1}:E_{PE2}:E_{PE3} \dots E_{PEn} = 1:\frac{1}{4}:\frac{1}{9} \dots \frac{1}{n^2}$
Total energy	$E = -\frac{2\pi^2Z^2k^2me^4}{n^2h^2}$; $E \propto \frac{Z^2}{n^2}$	$E_1:E_2:E_3 \dots E_n = 1:\frac{1}{4}:\frac{1}{9} \dots \frac{1}{n^2}$

Some Basic definitions

Excitation energy

The minimum energy required to excite an atom in the ground state to one of the higher stationary states is called excitation energy. (From $n = 1$ to $n = 1 \dots \infty$)

Consider the case of hydrogen atom.
We know that $E_1 = -13.6\ eV$ (Ground state),
$E_2 = -3.4eV$ (First Excited state),
$E_3 = -1.51eV$ (second excited state) and $E_n = \infty$.
In order to lift electron from ground state ($n = 1$) to the first excited state ($n = 2$) the energy required is:

$$E = E_2 - E_1 = -3.4 - (-13.6) = 10.2\ eV$$

So, in order to excite the atom from n=1 to n=2 we have to give the energy of 10.2 eV to Bombarded electrons used by this transition.

Excitation potential

The minimum accelerating potential which provides an electron energy sufficient to jump from the ground state (n=1) to one of the outer orbits is called excitation potential.

In order to lift electron from ground state ($n = 1$) to the first excited state ($n = 2$) the energy required is:

$$E = E_2 - E_1 = -3.4 - (-13.6) = 10.2\ eV$$

So, Excitation potential $= \frac{10.2eV}{e} = 10.2V$

It may be denoted that the value of excitation potential depends upon the state to which the atom is excited from the ground state.

Ionization energy

An atom has lost an electron is said to be ionized. Ionization energy is the energy necessary to remove an electron from the neutral atom. It is a minimum for the alkali metals which have a single electron outside a closed shell.

Consider the case of hydrogen atom, it has only one electron and this normally occupies the ground state. The energy of electron for $n = \infty$ state is zero and if

Ionization potential

The ionization potential is the potential through which an electron is accelerated for removal an electron from the neutral atom is called ionization potential.

Consider the case of hydrogen atom, it has only electron and this normally occupies the ground state. The energy of electron for $n = \infty$ state is zero and if

electron is level to this level ($n = \infty$), it becomes free of hydrogen atom. (hydrogen atom is ionized). $IE\ of\ H\ atom = E_{\infty} - E_1 = 0 - (-13.6) = 13.6eV$	electron is level to this level ($n = \infty$), it becomes free of hydrogen atom. (hydrogen atom is ionized). $IE\ of\ H\ atom = E_{\infty} - E_1 = 0 - (-13.6) = 13.6eV$ As the I.E of H atom = 13.6 eV $Ionization\ potential = \frac{13.6eV}{e} = 13.6V$

X-Ray

⇒ X-radiation is a form of electromagnetic radiation.

⇒ These are the radiation of the greater frequency or energy and lesser wavelength, Most X-rays have a wavelength ranging from 0.01 to 10 nanometers, corresponding to frequencies in the range 30 petahertz to 30 exahertz (3×10^{16} Hz to 3×10^{19} Hz) and energies in the range 100 eV to 100 keV.

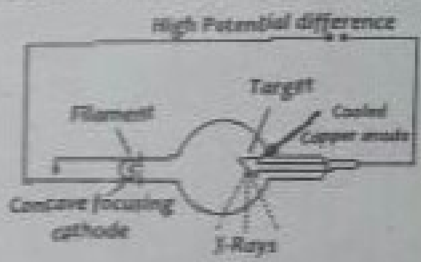
⇒ ***Why X-Rays are called 'X' Rays? and who discovered it?***

X-radiation is also referred to as Rontgen radiation, after Wilhelm Rontgen, who is usually credited as its discoverer, and who had named it X-radiation to signify an unknown type of radiation produced when electron collided with the walls of the tube.,

⇒ ***Production of X-rays:***

he modern X-ray tube, called Coolidge tube, is shown in the A heated element emits electrons which are accelerated towards a cooled copper anode under a high potential difference. A target metal of high atomic number and high melting point is lodged on the anode.

The intensity of the X-ray beam is controlled by the filament current controls the intensity of the X-ray beam by regulating the number of electrons striking the target per unit time. The potential difference between the cathode and anode controls the penetrating power of the beam.



⇒ ***Types of X-rays:***

Hard X – rays	Soft X – rays
High penetration power	Low penetration power
High frequency	Lower frequency
Lesser wavelength	Greater wavelength

Properties of X-rays:

✓ The wavelength of X-rays is very small in comparison to the wavelength of light. Hence they carry much more energy (This is the only difference between X-rays and light).
✓ They show all the important properties of light rays like; reflection, refraction, interference, diffraction and polarization etc
✓ used in the study of crystal structure.
✓ They have the ability to ionize gases.
✓ X-rays do not pass through heavy metals and bones.
✓ They affect photographic plates.
✓ They are invisible. (because they don't lie in visible region.
✓ Travel in a straight line.
✓ They are electromagnetic waves so it travels with the speed of light.
✓ X-rays are measure in terms of its unit of Rontgen.
✓ Chargeless.
✓ $\lambda_{gamma-rays} < \lambda_{x-rays} < \lambda_{UV-rays}$
✓ They produce photoelectric effect and Compton effect
✓ X-rays are not emitted by hydrogen atom.
✓ These cannot be used in Radar because they are not reflected by the target.
✓ Long exposure to X-rays is injurious for human body.
✓ Lead is the best absorber of X-rays.

Some important points and tricks

1) An atom is a stable entity.
2) Bohr 's model of the atom is only applicable for hydrogen atom.
3) In the Bohr model of the atom, an electron can revolve around the nucleus in a stable orbit without emitLing radiations if its orbit has whole number of de Broglie waves.
4) As the quantum number increases, the difference of energies between the consecutive energy levels decreases.
5) When an electron absorbs a photon, it is excited to the higher energy levels.
6) The binding energy of the electron in the ground state of hydrogen is called rycberg. 1 rydberg = 13.6 eV.
7) Balmer series was the first to be discovered in the hydrogen spectrum.
8) Only Lyman series is obtained by both emission as well as absorption. All other spectra are obtained only in the emission spectrum.
9) In an energy level diagram, the arrows pointing down represent emission lines while the arrows pointing up represent absorption lines.
10) In the hydrogen spectrum: Lyman series is in the ultraviolet region, Balmer series is in the visible region Paschen, Brackett and Pfund series are in the infrared region.
11) Observation of faint spectral line very near to blue line of hydrogen spectrum led to the discovery to deuterium.
12) Speed of electron in the ground state of the hydrogen atom is c/137 where, c = speed of light in vacuum.
13) Rydberg constant is different for different elements.
14) For an electron revolving in first orbit, the radius (r_1) of an orbit is $r_1 - 0.5\,A^o$. Thus, radius of n^{th} orbit is, $r_n = r_1 n^2$.
15) Angular frequency of electron in nth orbit

$$\omega_n \propto \frac{Z^2 m}{n^3} \propto \frac{1}{n^3} \quad \textit{for hydrogen atom.}$$

16) Energy of an electron in first orbit is $E_1 = -13.6$ eV. Thus, energy of an electron in nth orbit is, $E_n = \frac{E_1}{n^2}$

FORMULAS SHEET

1.	*Lyman series*	$R[\frac{1}{1^2} - \frac{1}{n_2^2}]$ *Value of* n_2 2,3,4,5,...
2.	*Balmer series*	$R[\frac{1}{2^2} - \frac{1}{n_2^2}]$ *Value of* n_2 3,4,5,6,...
3.	*Paschen series*	$R[\frac{1}{3^2} - \frac{1}{n_2^2}]$ *Value of* n_2 4,5,6,7,...
4.	*Bracket series*	$R[\frac{1}{4^2} - \frac{1}{n_2^2}]$ *Value of* n_2 5,6,7,8,...
5.	*Pfund series*	$R[\frac{1}{5^2} - \frac{1}{n_2^2}]$ *Value of* n_2 6,7,8,9,...
6.	General formula for any spectral series	$\frac{1}{\lambda} = R_H[\frac{1}{p^2} - \frac{1}{n^2}]$ $n > p$ where $n = p+1, p+2, p+3, ...$
7.	Angular momentum of electron in H-atom	$mvr = \frac{nh}{2\pi}$
8.	Energy of emitted photon	$hf = E_n - E_p$

✓ For X-ray photography of human body parts, $BaSO_4$ is the best absorber.

Classification of X-Rays

In X-ray tube, when high speed electrons strike the target, they penetrate the target. They lose their kinetic energy and comes to rest inside the metal. The electron before finally being stopped makes several collisions with the atoms in the target.

At each collision one of the following two types of X-rays may get formed.

(a) Characteristics x-rays

(b) Continuous x-rays

Characteristics X – rays

✓ Few of the fast-moving electrons having high velocity penetrate the surface atoms of the target material and knock out the tightly bound electrons even from the inner most shells of the atom. Now when the electron is knocked out, a vacancy is created at that place. To fill this vacancy electrons from higher shells jump to fill the created vacancies, we know that when an electron jumps from a higher energy orbit $E1$ to lower energy orbit E_2, it radiates energy $(E_1 - E_2)$. Thus this energy difference is radiated in the form of X-rays of very small but definite wavelength which depends upon the target material. The X-ray spectrum consists of sharp lines and is called characteristic X-ray spectrum.

Continuous X – rays

As an electron passes close to the positive nucleus of atom of the target, the electron is deflected from it's path as shown in figure.

This results in deceleration of the electron.

The loss in energy of the electron during deceleration is emitted in the form of X-rays.

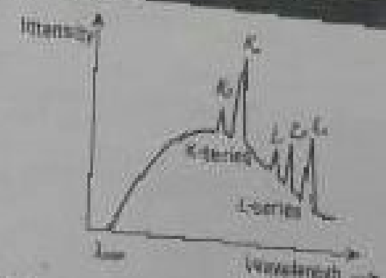
The X-ray photons so emitted form the continuous X-ray spectrum.

The wavelength at which the intensity is maximum depends on the accelerating voltage, being shorter for higher voltage and vice-versa.

$$\lambda \text{ at which intensity is max} \propto \frac{1}{Voltage}$$

Types of Characteristics X-rays and its intensity:

⇒ If the electron striking the target eject an electron from the K-shell of the atom, a vacancy is created in the K-shell. Immediately an electron from one of the outer shells, say L-shell jumps to the K-shell, emitting an X-ray photon of energy equal to the energy difference between the two shells. Similarly, if an electron from the M-shell jumps to the K-shell, X-ray photon of higher energy is emitted. The X-ray photons emitted due to the jump of electron from the L, M, N shells to the K-shells gives $K\alpha$, $K\beta$, $K\gamma$ lines of the K-series of the spectrum.



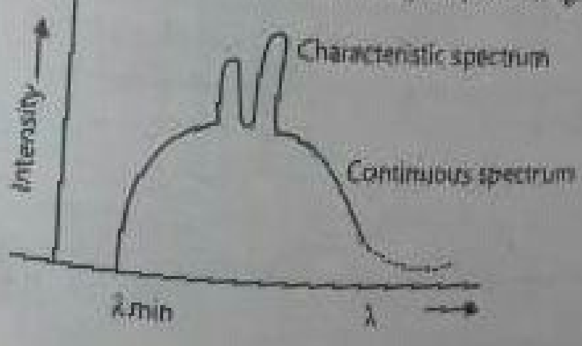
⇒ If the electron striking the target ejects an electron from the L-shell of the target atom, an electron from the M, N .... shells jump to the L-shell so that X-rays photons of lesser energy are emitted.

⇒ These photons form the L-series of the spectrum. In a similar way the formation of M series, N series etc. may be explained.

⇒ Intensity-wavelength graph: At certain sharply defined wavelengths, the intensity of X-rays is very large as marked $K\alpha$, $K\beta$ .... as shown in figure. These X-rays are known as characteristic X-rays. At other wavelengths the intensity varies gradually and these X-rays are called continuous X-rays.

X-Ray spectra:

A typical X-ray spectrum given by a target is shown in the Fig. The spectrum is basically continuous range of wavelengths starting from a minimum value. A line spectrum having sharp wavelengths is superimposed on this



1.	Velocity of electron in H-atom	$v_n = \frac{nh}{2\pi m r_n} \rightarrow v_n = \frac{2\pi k e^2}{nh}$
2.	Radius of nth orbit in hydrogen atom	$r = \frac{n^2 h^2}{4\pi^2 m k Z e^2} \rightarrow r_n = n^2 \times 0.53\ \text{Å}$
3.	KE of electron in nth orbit	$KE = \frac{1}{2} m v_n^2 \rightarrow KE = \frac{Ke^2}{2r_n}$
4.	PE of electron in nth orbit	$U = -q\Delta V \rightarrow U = -\frac{ke^2}{r}$
5.	TE of electron in nth orbit	$E_n = -\left(\frac{2\pi^2 k^2 m e^4}{h^2}\right)\frac{1}{n^2} \rightarrow E_n = -\frac{E_o}{n^2} \rightarrow -\frac{13.6}{n^2}\ eV$
6.	Emission spectra of hydrogen atom	$\frac{1}{\lambda} = \frac{E_o}{hc}\left(\frac{1}{p^2} - \frac{1}{n^2}\right) \rightarrow R_H\left(\frac{1}{p^2} - \frac{1}{n^2}\right)$
7.	k_α characteristics x-rays	$hf_{K\alpha} = E_L - E_K$
8.	k_β characteristics x-rays	$hf_{K\beta} = E_M - E_K$
9.	Maximum energy of breaking radiation	$KE = hf_{max} = \frac{hc}{\lambda_{min}}$
10.	Minimum wavelength of braking radiations	$\lambda_{min} = \frac{hc}{KE} = \frac{hc}{Ve}$

UNITS

6.	Rydberg's constant	m^{-1}	
7.	Planck's constant	$J.s$	
8.	Energy of electron in an orbit	J	eV
9.	Speed of light	m/s	
10.	Angular momentum	$J-s$	

CONSTANTS

5.	Rydberg constant	$1.097 \times 10^{7}\ m^{-1}$
6.	Planck's constant	$6.63 \times 10^{-34}\ J.s$
7.	Speed of light	$3 \times 10^{8}\ m/s$

MULTIPLE CHOICE QUESTIONS

...ording to classical theory of Rutherford model, the path of electron will be
...liptic
...yperbolic
...ircular
...piral

...r's atom model assumes
...e nucleus is of infinite mass and is at rest.
...lectron in a quantized orbit will not radiate energy.
...ss of electron remains constant.
...l the above conditions.

...r's model of hydrogen atom, which of the following pairs of quantities are quantized?
...ergy and linear momentum.
...ear and angular momentum.
...ergy and angular momentum.
...ergy but not the angular momentum.

4) The concept of stationary orbits was proposed by
a) Niels Bohr
b) J.J. Thomson
c) Rutherford
d) I. Newton

5) The ratio of the velocity of the electron in the first orbit to that in the second orbit is
a) 8:1
b) 4:1
c) 2:1
d) 1:1

6) In Bohr's model of hydrogen atom, the period of revolution of the electron in any orbit is proportional to
a) the quantum numbers.
b) square of the quantum number.
c) square root of the quantum number.
d) cube of the quantum number.

7) In the following transitions, which one has higher frequency?
a) 3 − 2
b) 4 − 3

c) 4 − 2 d) 3 − 1

8) When an electron in Hydrogen atom revolves in stationary orbit, it
a) does not radiate light though its velocity changes.
b) does not radiate light and velocity remains unchanged.
c) radiates light but its velocity is unchanged.
d) radiates light with the change of energy.

9) Energy of electron in a orbit of H-atom is
a) positive. b) negative.
c) zero. d) nothing can be said

10) In Bohr's model of hydrogen atom, let P.E. represent potential energy and T.E. the total energy. In going to a higher level
a) P.E. decreases, T.E. increases.
b) P.E. increases, T.E. increases.
c) P.E. decreases, T.E. decreases.
d) P.E. increases, T.E. decreases.

11) The kinetic energy of an electron revolving around a nucleus will be
a) four times of P.E. b) double of P.E.
c) equal to P.E. d) half of its P.E.

12) When a hydrogen atom is raised from the ground state to an excited state.
a) P.E. increases and K.E. decreases.
b) P.E. decreases and K.E. increases.
c) Both kinetic energy and potential energy increase.
d) Both K.E. and P.E. decrease.

13) In Bohr model of the hydrogen atom, the lowest orbit (nearest to nucleus) corresponds to
a) infinite energy.
b) the maximum energy.
c) the minimum energy.
d) zero energy.

14) The ratio of the angular momentum of the orbital electron in the first orbit to that in the 2nd orbit is
a) 2 : 1 b) 1 : 1
c) 1 : 2 d) none of these

15) When electron jumps from n = 4 level to n = 1 level, the angular momentum of electron changes by
a) $h/2\pi$
b) $\frac{2h}{2\pi}$
c) $3h/2\pi$
d) $4h/2\pi$

16) The radius of electron's second stationary orbit in Bohr's atom is R. The radius of the third orbit will be
a) 3R b) 2.25R
c) 9R d) R/3

17) Energy of an electron at infinity from nucleus is______
a) infinite b) minimum positive
c) maximum negative d) zero

18) To accelerate an electron so that it acquires energy to jump from present orbit to infinite orbit, electron needs______
a) ionization potential b) excitation potential
c) ionization energy d) excitation energy

19) Energy given externally is absorbed by an electron, then
a) electron jumps from present orbit to infinite orbit.
b) electron jumps from upper orbit to lower orbit.
c) electron is excited from ground state to upper state.
d) electron remains in same state

20) The ratio of the energies of the hydrogen atom in its first to second excited state is
a) ¼
b) 4/9
c) 9/4
d) 4

21) If the energy of a hydrogen atom in nth orbit is E_n, then energy in nth orbit of a singly ionized helium atom will be
a) $4E_n$
b) $\frac{E_n}{4}$
c) $2E_n$
d) $\frac{E_n}{2}$

22) In which of the following system will be radius of the first orbit (n = 1) be minimum
a) Deuterium atom
b) hydrogen atom
c) Doubly ionized lithium
d) Singly ionized helium

23) Which of the following element has highest first ionization potential
a) As b) Bi
c) Sb d) N

24) White light is passed through a dilute solution of potassium permanganate. The spectrum produced by the emergent light is
a) band emission spectrum
b) line emission spectrum
c) band absorption spectrum
d) line absorption spectrum

25) According to Bohr's theory the radius of electron in an orbit described by Principal quantum number n and atomic number Z is Proportional to
a) Z^2n^2 b) n^2/Z
c) Z^2/n d) $\frac{n^2}{Z^2}$

26) Number of spectral lines in hydrogen atom is
a) 6 b) 8
c) 15 d) infinite

27) Which of the following series in the spectrum of hydrogen alone lies in the visible region of the electromagnetic spectrum?
a) Paschen b) Lyman
c) Bracket d) Balmer

28) which of the following transition in hydrogen atoms emits Photon of highest frequency?
a) $n = 2$ to $n = 6$
b) $n = 1$ to $n = 2$
c) $n = 2$ to $n = 1$
d) $n = 6$ to $n = 2$

29) The innermost orbit of the hydrogen atom has a radius 0.53 Å. what is radius of 2nd orbit is?
a) 2.12 Å b) 1.06 Å
c) 21.2 Å d) 10.6 Å

30) The ratio of minimum to maximum wave length in Balmer series is
a) 1/4 b) 5/36
c) 3/4 d) 5/9

31) In terms of Rydberg constant R. The wave number of first Balmer line is
a) 5R/36 b) 8R/9
c) R d) 8R/20

32) The radius of hydrogen atom in the first excited level is
a) Twice b) four times
c) same d) half

33) The velocity of an electron in the inner most orbit of an atom is
a) highest b) lowest
c) mean d) zero

34) The X-rays cannot be diffracted by means of an ordinary grating because of
a) high speed b) short wave length
c) large wave length d) none of these

35) The spectrum obtained from a sodium vapor lamp is an example of
a) band spectrum b) continuous spectrum
c) emission spectrum d) absorption spectrum

36) Out of the following which is not a possible energy for photon to be emitted by hydrogen atom according to Bohr's atomic model
a) 1.9 eV b) 11.1eV
c) 13.6eV d) 0.65eV

37) The Bohr model of atoms
a) Predicts the same emission spectra for all types of atoms
b) assumes that the angular momentum of electrons is quantized
c) uses Einstein's photoelectric equation
d) predicts continuous emission spectra for atoms

38) The total energy of an electron in the first excited state of hydrogen atom is about -3.4eV. Its kinetic energy in this state is
a) 3.4 eV b) 6.8eV
c) -3.4eV d) -6.8eV

39) When hydrogen atom is in its first excited level its radius is
a) four times its ground state radius
b) twice
c) same d) half

40) In terms of Bohr's radius a_0, the radius of the second Bohr orbit of Hydrogen atom is given by
a) $4a_0$ b) $8a_0$
c) $\sqrt{2} \times a_0$ d) $2a_0$

41) When electron jump from the fourth orbit to the second orbit, one gets the
a) second line of Lyman series
b) second line of Paschen series
c) second line of Balmer series
d) first line of Pfund series

42) The wavelength of the first line of Lyman series for hydrogen atom is equal to that of the second line of Balmer series for a hydrogen like ion. The atomic number of Z of hydrogen like ion is
a) 3 b) 4
c) 1 d) 2

43) Which of the following transitions in hydrogen atoms emits the photon of highest frequency
a) n=2 to n=1 b) n=2 to n=6
c) n=6 to n=2 d) n=1 to n=2

44) the energy of hydrogen atom in nth orbit is En then the energy in the nth orbit of a single ionized helium atom will be
a) 4En
b) En/4
c) 2En
d) En/2

45) The ionization energy of the electron in the hydrogen atom in its ground state is 13.6 eV. The atoms are excited to higher energy levels to emit radiations of 6 wavelengths. Maximum wave length of emitted radiation corresponds to the transition between...
a) $n = 3$ to $n = 1$ states
b) $n = 2$ to $n = 1$ states
c) $n = 4$ to $n = 3$ states
d) $n = 3$ to $n = 2$ states

46) In which of the following systems will the radius of the first orbit (n=1) be minimum
a) Hydrogen atom
b) Doubly ionized lithium
c) Singly ionized helium
d) Deuterium atom

1) According to Rutherford model, an atom have a positively charged nucleus and electrons revolve around the nucleus in circular path.
Answer: (c)

2) d
3) C
4) [illegible]
5) $v \propto \frac{1}{n}$

$$\frac{v_1}{v_2} = \frac{n_2}{n_1} = \frac{2}{1}$$

6) $v = \frac{2\pi r}{T} \rightarrow T = \frac{2\pi r}{v}$

$$\text{but } r \propto n^2 \text{ and } v \propto \frac{1}{n} \therefore T \propto \frac{r}{v} \propto \frac{n^2}{\left(\frac{1}{n}\right)} \propto n^3$$

7) 3-1 transition has higher energy.

$$\propto f = \frac{E}{h} \text{ so it has higher frequency}$$

8) [illegible]

9) The electron in the orbit is bound to nucleus. Hence, its total energy is taken negative.

10) $K.E \propto \frac{1}{n^2}$ and $T.E \propto \frac{1}{n^2}$

11) $P.E = -2K.E$

12) $P.E \propto -\frac{1}{r}$ and $K.E \propto \frac{1}{r}$

As r increases, K.E. decreases but P.E increases.

13) In hydrogen atom, the lowest orbit (n = 1) corresponds to minimum energy (- 13.6 eV).

14) Angular momentum $= mvr = \frac{nh}{2\pi}$

$$\text{Angular momentum} \propto n$$

$$\text{Ratio} \rightarrow \frac{n_1}{n_2} = \frac{1}{2}$$

15) Angular momentum, $L = \frac{nh}{2\pi}$

$$\Delta L = L_4 - L_1 = \frac{4h}{2\pi} - \frac{h}{2\pi} = \frac{3h}{2\pi}$$

16) $r_n \propto n^2 \Rightarrow \frac{r_3}{r_2} = \frac{9}{4} \Rightarrow r_3 = \frac{9}{4}R = 2.25\,R$

17) [illegible]
18) [illegible]
19) C

20) First excited state i.e., second orbit (n = 2). Second excited state i.e., third orbit (n = 3).

$$\frac{r_2}{r_3} = \frac{E_3}{E_2} = \left(\frac{2}{3}\right)^2 = \frac{4}{9}$$

21) Energy of helium atom in nth orbit $= E_n$. Energy of the singly ionized helium atom is 4 times that of hydrogen atom. Energy of singly ionized helium atom $= 4\,E_1$
Answer: (a)

22) Radius of the orbit is inversely proportional to the nuclear charge (Ze) which is maximum in [illegible] doubly ionized lithium the radius of the first [illegible] is minimum.
Answer: (c)

23) The valence electrons in nitrogen are bound by a stronger force than that in the case of As, Sb and Bi. Thus, first ionization potential of nitrogen is the highest among the given elements.
Answer: (d)

24) Answer: (c)

25) Bohr's radius $= r = \frac{n^2 h^2 \varepsilon_0}{\pi Z e^2 m} \Rightarrow r \propto \frac{n^2}{Z}$
Answer: (b)

26) Answer: (d)

27) Answer: (d)

28) n=2 to n=6 absorbs photon, n=9 to n=2 absorbs photon, n=6 to n=2 emission of photon, n=2 to n=1 emission of photon.

$$\Delta E = E_2 - E_6 = \frac{-13.6}{36} - \frac{-13.6}{4} = 3.02\ eV$$

$$\Delta E' = E_2 - E_1 = \frac{-13.6}{4} - \frac{-13.6}{1} = 10.2\ eV$$

$$\therefore \Delta E' > \Delta E \text{ and hence } hf' > hf$$

Answer: (c)

29) $r \propto \frac{n^2}{Z}$ for hydrogen atom $Z = 1$

$r \propto n^2 \Rightarrow \frac{r_1}{r_2} = \frac{1}{4} \Rightarrow r_2 = 4r_1 = 4 \times 0.53 =$ 2.12
Answer: (a)

30) $\frac{1}{\lambda} = R\left[\frac{1}{3^2} - \frac{1}{n^2}\right] \rightarrow \lambda_1 = \frac{9}{R}$

$$\text{putting } n = \infty$$

$$\lambda_2 = \frac{36}{5R} \rightarrow \frac{\lambda_1}{\lambda_2} = \frac{4}{R} \times \frac{5R}{36} = \frac{5}{9}$$

Answer: (d)

31) $\text{wavenumber} = \frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{3^2}\right]$

$$\frac{1}{\lambda} = R\left[\frac{1}{4} - \frac{1}{9}\right] = \frac{5R}{36}$$

Answer: (a)

32) Radius of H-atom varies as a square of n. So for excitation from n=1 to n=2, radius becomes 4 times.
Answer: (b)

33) As the radius decreases orbital speed increases then option (a) is correct.
Answer: (a)

34) We know that the X-rays are of short-wave length as compared to grating constant of optical grating. As a result of this, it makes difficult to observe X-rays diffraction with ordinary grating.
Answer: (b)

35) A spectrum is observed, when light coming directly from a source is examined with a spectroscope. Therefore, spectrum obtained from a sodium vapor lamp is emission spectrum.
Answer: (c)

36) Energy for orbit is given by $E = -\frac{13.6}{n^2}$

$$n = 1\ E_1 = -13.6\ eV$$

$$n = 2\ E_2 = -3.4\ eV$$

$$n = 3\ E_3 = -1.51\ eV$$

$$n = 4\ E_4 = -0.85\ eV$$

$n = 5\ E_5 = -0.58eV$

Thus difference of 11.1 eV is not possible

Answer: (b)

37) In Bohr's model angular momentum is quantize angular momentum=$nh/2\pi$

Answer:(b)

38) Kinetic energy = - Mechanical energy

Mechanical energy= $-\frac{13.6}{n^2} = -\frac{13.6}{2^2}$

For first excited state n=2

Mechanical energy=-13.6/4=-3.4eV

Thus kinetic energy=+3.4 eV

Answer: (a)

39) radius of hydrogen atom R∝ n2

now for first excited state n=2. thus radius of first excited state is four times the radius in ground state.

Answer: (a)

40) radius of hydrogen atom R∝ n^2

∴ radius of 2nd Bohr's orbit=$4a_0$

Answer:(a)

41) When the electron drops from any orbit to second orbit, then wavelength of line obtained belong to Balmer series.

Answer: (c)

42) According to equation for spectral line emitted by hydrogen like ion is given by

$$\frac{1}{\lambda} = Z^2\left[\frac{1}{n^2} - \frac{1}{n_2^2}\right]$$

For Lyman series of hydrogen Z=1 thus

$$\frac{1}{\lambda_1} = R\left[\frac{1}{1^2} - \frac{1}{2^2}\right]$$

For second line of Balmer series of Hydrogen like ion

$$\frac{1}{\lambda_2} = RZ^2\left[\frac{1}{2^2} - \frac{1}{4^2}\right]$$

Given $\lambda_1 = \lambda_2$

On simplification Z=2

Answer: (d)

43) Frequency of wave emitted

$$f = R\left[\frac{1}{n^2} - \frac{1}{m^2}\right]$$

Transition from higher m to lower energy level should have least difference to have maximum frequency

in given case m=2 and n=1 will be the answer

Answer: (a)

44) Energy of nth orbit $En \propto \frac{Z^2}{n^2}$

For helium Z = 2

∴ Energy of nth orbit of Helium $E_n = 4E_n$

Answer:(a)

45) Number of wavelengths= $\frac{n(n-1)}{2}$

Here n is the energy state

$$\Rightarrow 6 = \frac{n(n-1)}{2}$$

n=4 thus answer is n=4 to n=3

Answer: (c)

46) radius $R \propto \frac{1}{Z}$

For lithium $Z = 3$ thus Li^{+2}

Answer: (b)

47) Energy of ground state is 13.6 eV and excited by radiation of 12.1eV. Thus, excited state energy= $-13.6 + 12.1 = 1.5\ eV$

$$En = -\frac{13.6}{n^2}$$

From above formula, Energy of third energy level is -1.5

Thus possible transitions are 1→2, 1→3, 2→3 so three lines are possible

Answer: (a)

48) Answer: (c)

49) $P.E = -\frac{KZe^2}{r}$

$$K.E = \frac{KZe^2}{2r}$$

Here 'r' is the radius of the orbit which increases as we move from ground to an excited state. Therefore, when a hydrogen atom is raised from the ground state, it increases the value of 'r'. As a result of this P.E. increases (decrease in negative) and K.E. decreases

Answer: (b)

50) 53 electrons in iodine atom are distributed as 2,8,18,7

$\Rightarrow n = 5$ and $Z = 53$

formula for radius is

$$r_n = (0.53 \times 10^{-10})\frac{n^2}{Z}$$

$$r_5 = (0.53 \times 10^{-10})\frac{5^2}{53} = 2.5 \times 10^{-11}m$$

Answer:(a)

51) $E \propto Z^2$ for helium $Z = 2$

$\therefore E_{He} = 4 \times (2)^2$

$E_{He} = 2 \times 13.6 = 54.4eV$

Answer: (d)

52) From Bohr postulate $mvr = \frac{nh}{2\pi}$

$mv = \frac{nh}{2\pi r}$ - - - eq(1)

Now $\lambda = \frac{h}{mv}$ - - - eq(2)

From eq(1) and eq(2) we get $\lambda = (2\pi r/n)$

$\therefore 2\pi r = n\lambda$

∴ circumference = $n\lambda$

∴ n waves are permitted in nth orbit

Answer: (a)

53) Answer: (c)

54) Answer:(d)

55) Answer: (a)

56) Answer: (d)

57) Answer: (a)

58) Wave length of ray

$$\lambda = \frac{hc}{E} = 0.86\ A^0$$

Since λ < 100 Å

So it is X-ray

Answer:(b)

59) To produced X-ray electrons are bombarded on the target. Larger the number of electrons bombarded greater will be intensity of X-rays.
Answer: (d)

60) Infrared photons have frequency less than red light so its energy will be less than infrared light.
Answer:(a)

61) Bohr's theory was modified by Planck who gave quantum mechanical approach to structure of atom.
Answer: (b)

62) $Total\ energy = K.E. + P.E.$

$$3.555\ MeV = K.E. + 0$$
$$K.E. = 3.555\ MEV$$

Answer:(a)

63) For Lyman series

$\frac{1}{\lambda} = \left[\frac{1}{1^2} - \frac{1}{n^2}\right]$

$for\ shortest\ wavelength\ n = \infty$

$\frac{1}{\lambda_1} = R \rightarrow \lambda_1 = \frac{1}{R}$

64) $for\ balmer\ series\ \frac{1}{\lambda_2} = R\left[\frac{1}{2^2} - \frac{1}{n^2}\right]$

$for\ shortest\ wavelength\ n = \infty$

$\frac{1}{\lambda_2} = \frac{R}{4} \Rightarrow \lambda_2 = \frac{4}{R}$ $and\ the\ ratio\ \frac{\lambda_2}{\lambda_1} = \frac{4}{1}$

Answer:(a)

65) Rutherford's experiment disclosed that atoms are mostly hollow and, in the center, there is a small sized nucleus. Thus, it discovers the position of nucleus.
Answer: (c)

66) $E_2 = -\frac{E_1}{n^2} = -\frac{13.6}{2^2} = -3.4\ eV$
Answer:(d)